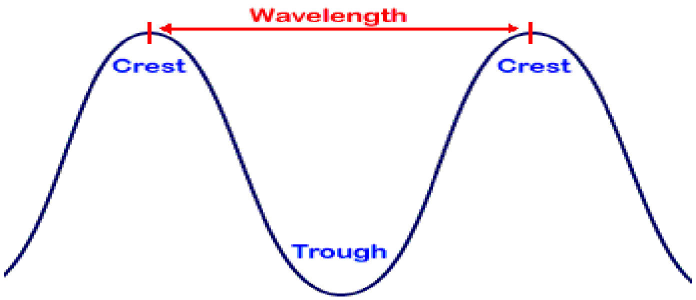
Electromagnetic Radiation

Properties of E-M Radiation

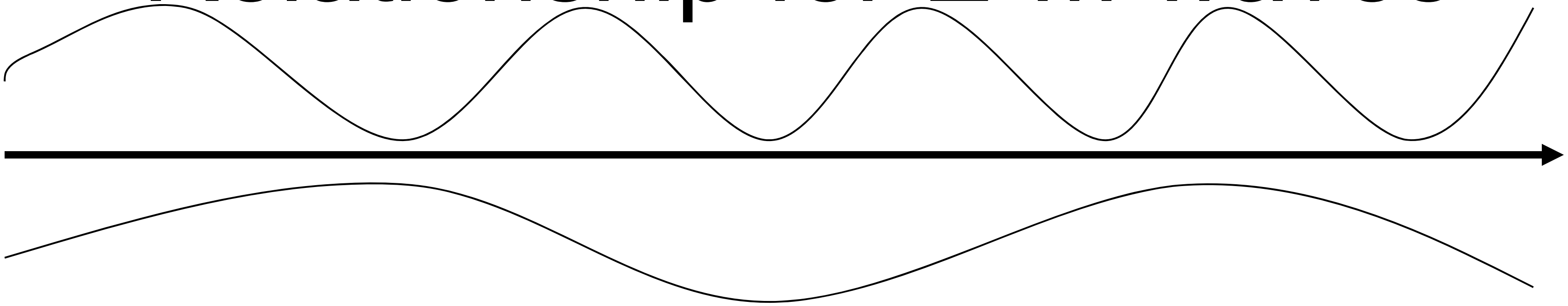
- All electromagnetic radiation has the following common properties:
  1. It is energy and has no mass.
  2. It travels at the speed of light.
  3. It can travel through a vacuum.
  4. It comes from atoms.
  5. It travels in a wave motion.
  6. It moves through space as bundles of energy called photons.
  7. Each wave has a frequency which is related to the energy. The higher the frequency, the higher the energy.

Wavelength and Frequency



- The wavelength is the distance from a point on a wave to when that point is repeated on the next wave.
- Measured in meters per cycle, but cycle is ignored so the units are meters.
- The frequency is how many cycles pass by a point in a second.
- Units are 1/sec or Hertz (Hz)

Relationship for E-M waves



- Since both waves travel at the same speed (the speed of light) then there is a relationship between frequency and wavelength.
- The shorter the wavelength (top) the higher the frequency.
- The longer the wavelength (bottom) the lower the frequency.
- This is an inverse relationship.

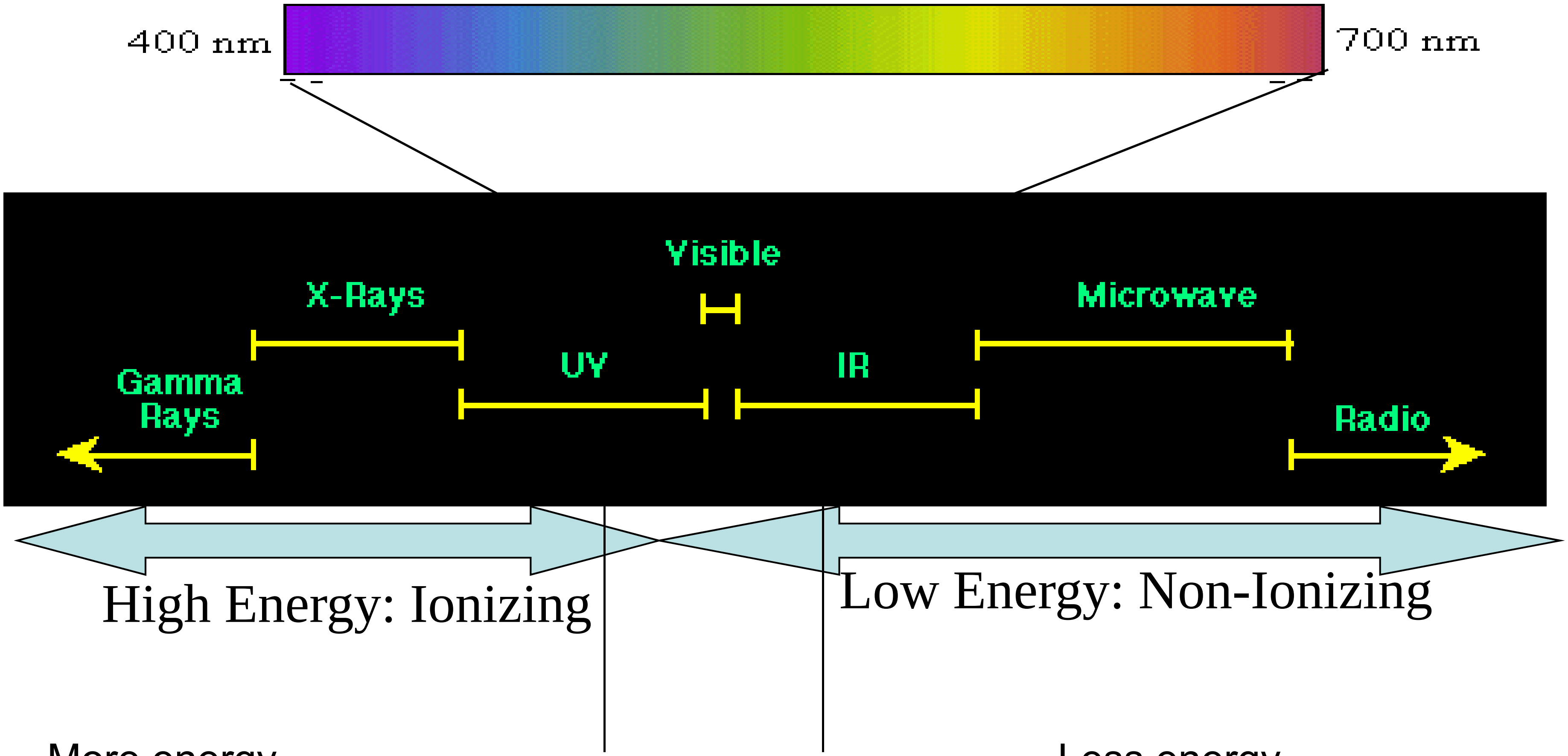
Equation

- The relationship is that the frequency times the wavelength is always equal to the speed of light.
- This value for the constant velocity was determined by Maxwell.
- $c = \lambda v$
- c = speed of light (3×10^8 m/s)
- λ = wavelength in meters
- v = frequency in 1/sec

Energy of E-M waves

- Energy is proportional to frequency.
- This was discovered by Max Planck, and the relationship is related to a quantity called Planck’s Constant.
- Planck’s Constant (h) is equal to
  6.626×10^{-34} Js
- The equation is $E = hv$
- Energy has units of Joules, and frequency is in 1/sec.

The Electromagnetic Spectrum



High Energy: Ionizing

Low Energy: Non-Ionizing

Sunlight

More energy	Less energy
High frequency	Low frequency
Short wavelength	Long wavelength

Particle view of E-M Radiation

- Einstein discovered that light can be described as a particle.
- He found that light traveled as little packets of energy.
- He called these photons, and proved their existence through an experiment about the photoelectric effect.
- This is where light of low energy cannot force electrons to be emitted from a metal no matter how much light you have.
- But even if you have very little high energy light, electrons will begin to be emitted.

Mass of E-M Radiation

- Since the electromagnetic radiation can be described like a particle, Einstein’s theory of relativity says it will behave as if it has a mass.
- This relationship is described by his famous $E=mc^2$ equation.
- E is the energy in Joules
- m is the mass in kilograms
- c is the speed of light.

Binding Energy

- It takes energy to hold particles together.
- Whenever the particles are split apart, this lost energy can be measured as a loss of mass.
- In nuclear fission, the lost strong force needed to hold the nucleus together equals the release of energy.
- Calculate the loss of mass and plug into $E=mc^2$ to determine the energy.
- This is nuclear binding energy.

Nuclear Mass Defect

- When the protons and neutrons in a nucleus are put together, their combined mass is less than when they are apart.
- This is nuclear mass defect, so you add up the mass of the protons and neutrons separately, and then subtract that from the mass of the nucleus combined.
- Plug the difference into $E=mc^2$ and solve for energy.
- This is the energy released in a fusion reaction.

The atom

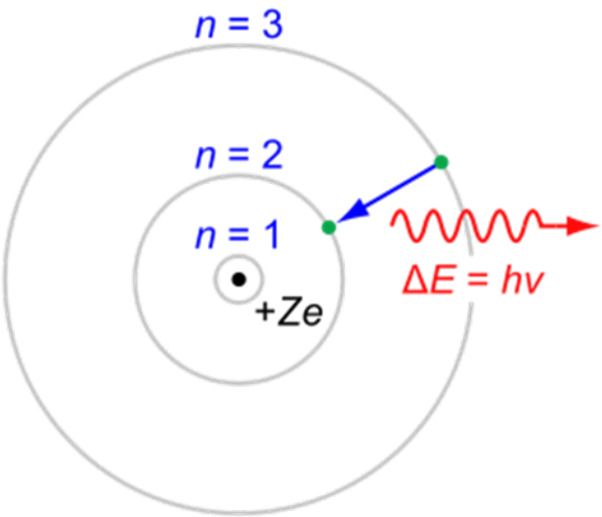
- Dalton had pictured an atom like a marble.
- Thomson saw it as a positive charged transparent object with the electrons imbedded in it.
- Rutherford’s gold foil experiment showed that there was a small nucleus with the electrons orbiting.
- There were issues though with Rutherford’s model and classical physics because the electrons would have to emit energy to stay in orbit, and that would mean they would crash into the nucleus and be lost.

Bohr

- Neils Bohr used Planck’s idea of quantized energy to explain how the electrons could orbit without losing energy.
- He stated that the electrons can only have specific amounts of energy in each energy level.
- The energy required to orbit in each path was very specific, and electrons would need exact amounts of energy to jump to higher energy levels.
- Any more or less energy, and the electron will not absorb it.

Energy levels

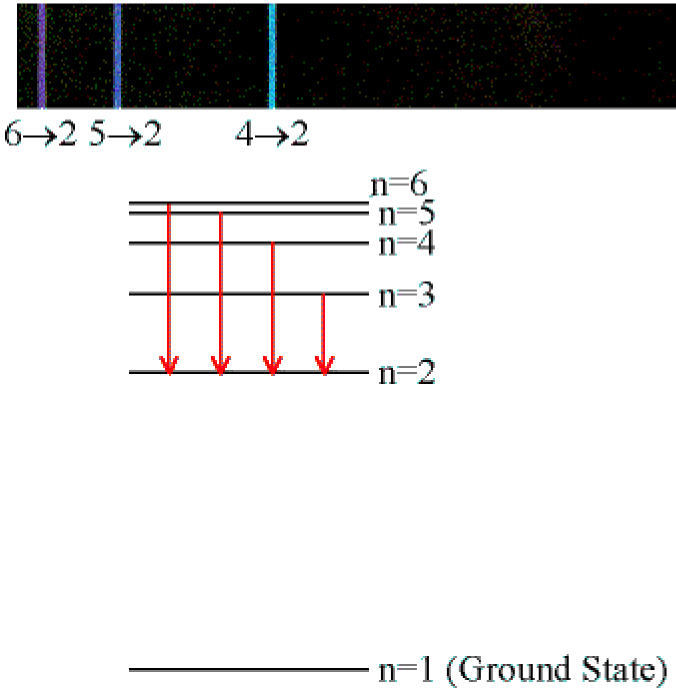
- Bohr stated that electrons have a “ground state”, or a minimum energy.
- For a hydrogen atom (which is what he worked with) n=1 was the ground state.
- If you gave it the right amount of energy it would jump to higher energy levels



When it dropped back to a lower energy level it released E-M radiation which had the energy equal to the difference in energy levels.

Balmer Series

- Bohr predicted mathematically what the wavelengths of visible light would be for a hydrogen atom.
- This was called the Balmer series since he was the one who first saw it and verified Bohr’s predictions.



Emission Spectra

- All elements have electrons that can gain or lose energy and change energy levels.
- As more electrons are added to an atom, the energy of each electron will be slightly different. This changes the energies needed to change energy levels.
- Each element then gives off different looking colors when their electrons get excited.
- This gives us the ability to identify what element is present by that emitted color.
- These are called emission spectra.

More Complicated

- Bohr's predictions were verified, but it could not explain all the results of experiments so the atom must be more complicated then that.
- DeBroglie came up with the concept of electrons having particle and wave properties just like light.
- This was verified through experiment so suddenly electrons could be described by waves.

Uncertainty

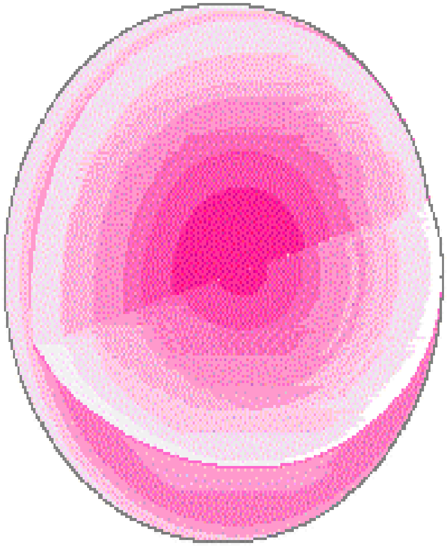
- Heisenberg realized that electrons were very small and moved very fast, and energy needed to hit them to tell us where they are.
- This meant that it was impossible to determine accurately both the location and momentum of the electrons.
- The more you know one, the less you know the other.
- This led to the idea that electrons just had probable locations in the atom.

Quantum Model

- Schrodinger mathematically calculates the wave functions (equations of the wave properties of the electron) to show where electrons can be found.
- These equations carve out a volume of space where the electron is “probably” found.
- The shape of the space is called an orbital.
- Since it is a probability, the electron may not actually be in that volume all of the time.
- Different electron energies give different orbital shapes and energy levels.

s-orbitals

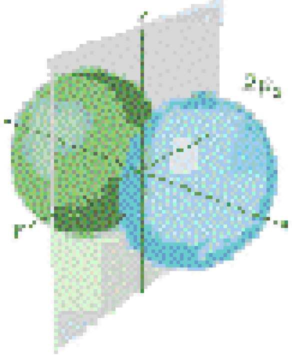
- Right now there are four orbital ty electrons may occupy.
- The simplest is the s- orbital which is a sphere.
- This is the first orbital that electrons fill at each energy level.

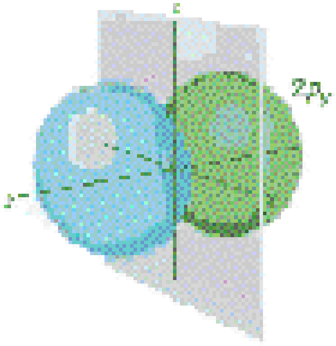


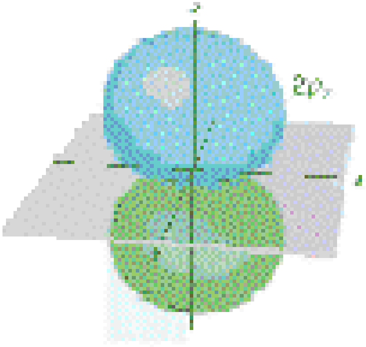
- Each higher energy level gives a large s-orbital.

p-orbitals

- p-orbitals are the second highest energy orbital in each energy level.
- The first p-orbitals appear in energy level 2.
- They look like 3-dimensional bowties.
- There are three different ways that they can be oriented, so there are 3 different p-orbitals.
- They get larger at higher energy levels too.







d and f orbitals

- In energy level 3 the first d-orbitals appear.
- In energy level 4 the first f-orbitals appear.
- Their shapes are very complex.
- There are 5 different d-orbitals.
- There are 7 different f-orbitals.
- All of the different orbitals or one type together is called a sublevel.
- So the p-sublevel contains 3 p-orbitals

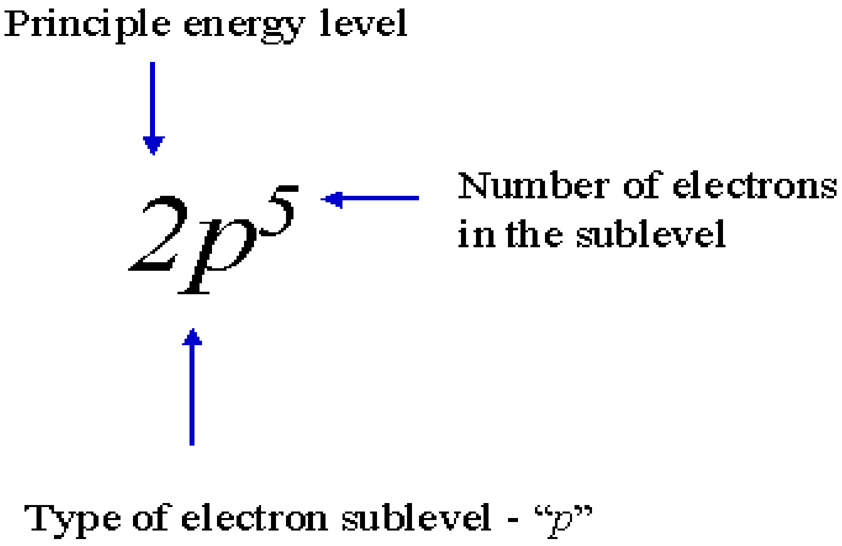
Rules of orbitals

- Aufbau – principle states that electrons when filling an atom will always choose the lowest energy orbital first.
- This doesn’t always follow the expected order since the energy that each electron has in an orbit is very complex.
- Pauli – his exclusion principle states that no more than two electrons with opposite spin can occupy the same orbital.
- This means an s sublevel can hold 2 electrons, a p can hold 6, a d can hold 10, and an f can hold 14.

Electron configurations

- This is a way of writing out the “address” of each electron in the atom.

Principal Energy Level	Number of sublevels	Name of Sublevels	Number of electrons	Type of sublevel
n=1	1	s	2	1*s*
n=2	2	s, p	8	2*s*, 2*p*
n=3	3	s,p,d	18	3*s* 3*p* 3*d*
n=4	4	s,p,d,f	32	4*s* 4*p* 4*d* 4*f*



Principle energy level

$2p^5$

Number of electrons in the sublevel

Type of electron sublevel - “*p*”

Seaborg’s Periodic Table

- Recognizing that orbitals determine where electrons go, and therefore the chemical properties of the elements, Seaborg recreated the order of the periodic table.
- This is the modern periodic table we have today.
- Groups 1 and 2 are the s-block and contain s-electrons. (2 columns)
- Groups 13-18 are the p-block. (6 columns)
- The transition elements are the d-block.
- The 2 rows at the bottom of the f block.

Electron Configurations of Elements.

- Use the periodic table as a guide.
- Each row represents the energy level.
- Even though the d-orbitals should appear in the third energy level, they actually come after the 4s.
- This is because d electrons are very high energy, and are more energy than the 4s and Aufbau states electrons will go to the lowest energy first.
- That’s why the transition elements don’t appear until row 4. They are always 1 energy level behind.
- The f electrons are 2 energy levels behind.